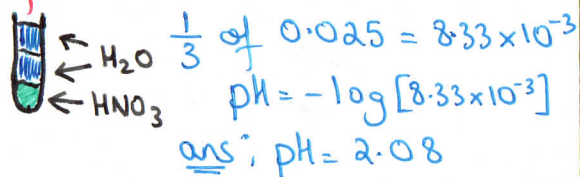


$$pH = -\log [H_3O^+]$$

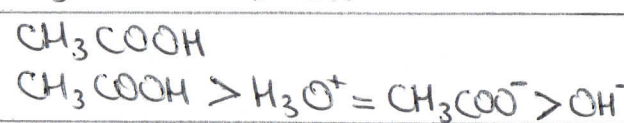
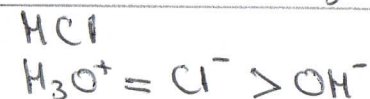
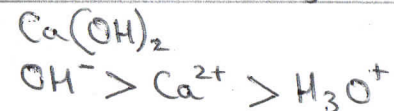
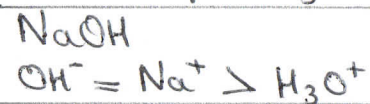
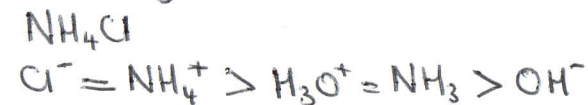
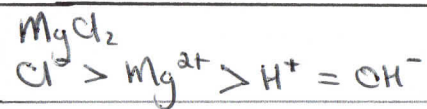
calculate pH of 10ml of 0.025M HNO_3 if 20ml of water added to it



$$K_w = 1 \times 10^{-14} = [OH^-] [H_3O^+]$$

calculate $[OH^-]$ of the soln

$$[OH^-] = \frac{1 \times 10^{-14}}{8.33 \times 10^{-3}} = 1.20 \times 10^{-12}$$



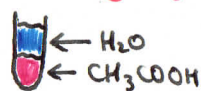
$$K_a = \frac{[H_3O^+] [A^-]}{[HA]}$$

acid dissociation constant

$$pK_a = -\log K_a$$

calculate the pH if 20ml of water is added to 20ml of 0.05M ethanoic acid

$$K_a(CH_3COOH) = 1.78 \times 10^{-5}$$



conc $\frac{0.05}{2} = 0.025M$



$$K_a = \frac{[CH_3COO^-] [H_3O^+]}{[CH_3COOH]}$$

$$1.78 \times 10^{-5} = \frac{x \cdot x}{0.025}$$

$$(1.78 \times 10^{-5}) \times (0.025) = x^2$$

$$6.67 \times 10^{-4} = x$$

$$pH = -\log [H_3O^+]$$

$$= -\log (6.67 \times 10^{-4})$$

$$pH = 3.18$$

K_a value describes the extent that the acid will react with water. CH_3COOH is a weak acid so the K_a value is low, indicating that it is only slightly reactive with water, so K_a must be used in the calculation

$$K_b = \frac{[BH^+] [OH^-]}{[B]}$$

$$K_b = \frac{K_w}{K_a}$$

calculate the pH of a 0.15M solution of NH_3

$$K_a(NH_4^+) = 5.75 \times 10^{-10}$$

$$K_b = \frac{K_w}{K_a} = \frac{(1 \times 10^{-14})}{(5.75 \times 10^{-10})}$$

$$K_b = 1.74 \times 10^{-5}$$

$$K_b = \frac{[NH_4^+] [OH^-]}{[NH_3]}$$

$$1.74 \times 10^{-5} = \frac{x \cdot x}{0.15}$$

$$(1.74 \times 10^{-5}) (0.15) = x^2$$

$$1.62 \times 10^{-3} = x$$

$$K_w = [OH^-] [H_3O^+]$$

$$\frac{1 \times 10^{-14}}{1.62 \times 10^{-3}} = [H_3O^+]$$

$$6.17 \times 10^{-12} = [H_3O^+]$$

$$pH = -\log (6.17 \times 10^{-12})$$

$$pH = 11.2$$